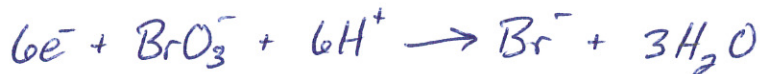


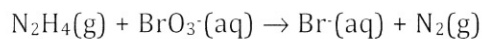
Unit 17 Practice Test: Electrochemistry

1. Under acidic conditions the bromate ion is reduced to the bromide ion. Write the balanced half-reaction for this process.

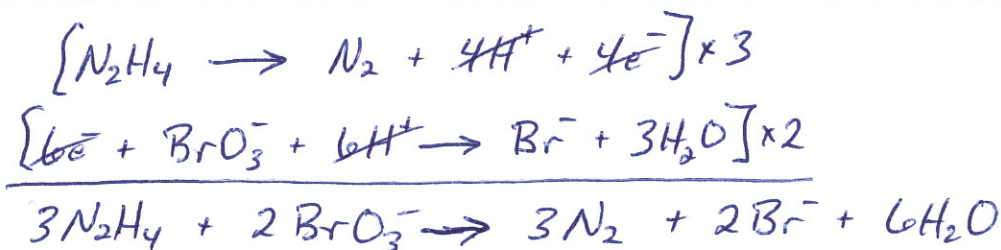
- a) $\text{BrO}_3^- + 6\text{H}^+ + 6\text{e}^- \rightarrow \text{Br}^- + 3\text{H}_2\text{O}$
 b) $2\text{BrO}_3^- + 6\text{H}^+ \rightarrow \text{Br}_2 + 6\text{H}_2\text{O} + 3\text{e}^-$
 c) $\text{BrO}_3^- + 6\text{H}_2\text{O} + 10\text{e}^- \rightarrow \text{Br}_2^- + 12\text{H}^+ + 3\text{O}_2$
 d) $2\text{BrO}_3^- + 6\text{H}_2\text{O} \rightarrow 2\text{Br}^- + 12\text{H}^+ + 6\text{O}_2 + 8\text{e}^-$
 e) $2\text{BrO}_3^- + 6\text{H}^+ \rightarrow \text{Br}_2^- + 3\text{H}_2\text{O} + 3\text{e}^-$



2. Balance the following redox equation which occurs in acidic solution.



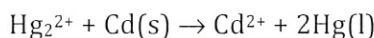
- a) $3\text{N}_2\text{H}_4 + \text{BrO}_3^- \rightarrow 3\text{N}_2 + \text{Br}^- + 3\text{H}_2\text{O} + 6\text{H}^+$
 b) $\text{N}_2\text{H}_4 + \text{BrO}_3^- + 2\text{H}^+ \rightarrow 2\text{Br}^- + \text{N}_2 + 3\text{H}_2\text{O}$
 c) $3\text{N}_2\text{H}_4 + 2\text{BrO}_3^- + 12\text{H}^+ \rightarrow 3\text{N}_2 + 2\text{Br}^- + 6\text{H}_2\text{O} + 12\text{H}^+$
 d) $\text{N}_2\text{H}_4 + 2\text{BrO}_3^- + 8\text{H}^+ \rightarrow 2\text{Br}^- + \text{N}_2 + 6\text{H}_2\text{O}$
 e) $3\text{N}_2\text{H}_4 + 2\text{BrO}_3^- \rightarrow 3\text{N}_2 + 2\text{Br}^- + 6\text{H}_2\text{O}$



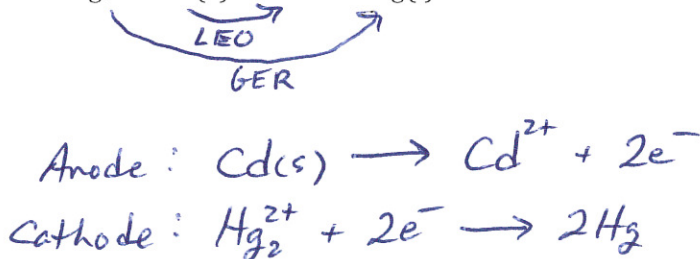
3. Which of the following reactions is NOT a redox reaction?

- a) $2\text{HgO}(\text{s}) \rightarrow 2\text{Hg}(\text{l}) + \text{O}_2(\text{g})$
 b) $\text{H}_2(\text{g}) + \text{Br}_2(\text{g}) \rightarrow 2\text{HBr}(\text{g})$
 c) $2\text{HCl}(\text{aq}) + \text{Zn}(\text{s}) \rightarrow \text{H}_2(\text{g}) + \text{ZnCl}_2(\text{aq})$ C stays 4+
 d) $\text{H}_2\text{CO}_3(\text{aq}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$ ← H stays 1+
 e) $2\text{KClO}_3 \rightarrow 2\text{KCl}(\text{s}) + 3\text{O}_2(\text{g})$ O stays 2-

4. Which of the following is the correct cell notation for the reaction



- a) $\text{Cd}^{2+} | \text{Cd} || \text{Hg}_2^{2+} | \text{Hg}$
 b) $\text{Cd}^{2+} | \text{Hg}_2^{2+} || \text{Cd} | \text{Hg}$
 c) $\text{Cd} | \text{Cd}^{2+} || \text{Hg}_2^{2+} | \text{Hg}$
 d) $\text{Cd}^{2+} | \text{Hg} || \text{Hg}_2^{2+} | \text{Cd}$
 e) $\text{Hg} | \text{Cd} || \text{Hg}_2^{2+} | \text{Cd}^{2+}$



Standard Reduction Potentials at 25°C E° (volts)

$F_2(g) + 2e^- \rightarrow 2F^-(aq)$	+2.87
$Au^{3+} + 3e^- \rightarrow Au(s)$	+1.50
$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$	+1.36
$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$	+1.23
$Br_2(l) + 2e^- \rightarrow 2Br^-(aq)$	+1.08
$Ag^+(aq) + e^- \rightarrow Ag(s)$	+0.80
$Hg_2^{2+}(aq) + 2e^- \rightarrow 2Hg(l)$	+0.79
$I_2(s) + 2e^- \rightarrow 2I^-(aq)$	+0.535
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	+0.337
$Sn^{4+}(aq) + 2e^- \rightarrow Sn^{2+}(aq)$	+0.15
$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$	-0.14
$Cd^{2+}(aq) + 2e^- \rightarrow Cd(s)$	-0.40
$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	-0.763
$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$	-0.828
$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$	-1.66
$K^+(aq) + e^- \rightarrow K(s)$	-2.93
$Li^+(aq) + e^- \rightarrow Li(s)$	-3.045

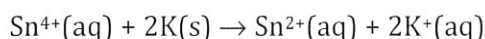
5. Given the two half reactions and their potentials, which net reaction is spontaneous?



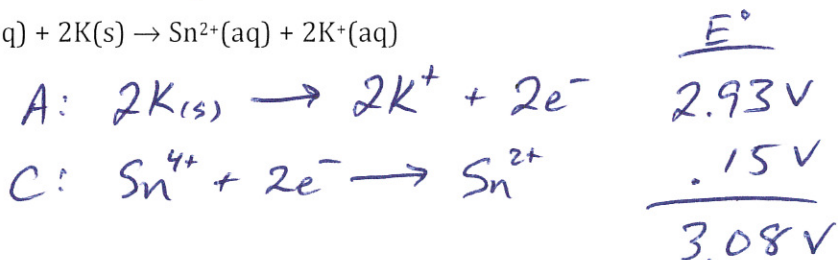
- a) $Ni(s) + Mg^{2+}(aq) \rightarrow Mg(s) + Ni^{2+}(aq)$
 b) $Ni^{2+}(aq) + Mg(s) \rightarrow Mg^{2+}(aq) + Ni(s)$
 c) $Ni(s) + Mg(s) \rightarrow Mg^{2+}(aq) + Ni^{2+}(aq)$
 d) $Mg^{2+}(aq) + Ni^{2+}(aq) \rightarrow Mg(s) + Ni(s)$
 e) $Mg^{2+}(aq) + Mg(s) \rightarrow Ni(s) + Ni^{2+}(aq)$



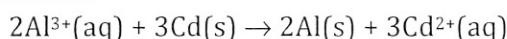
6. Calculate E° for the following reaction:



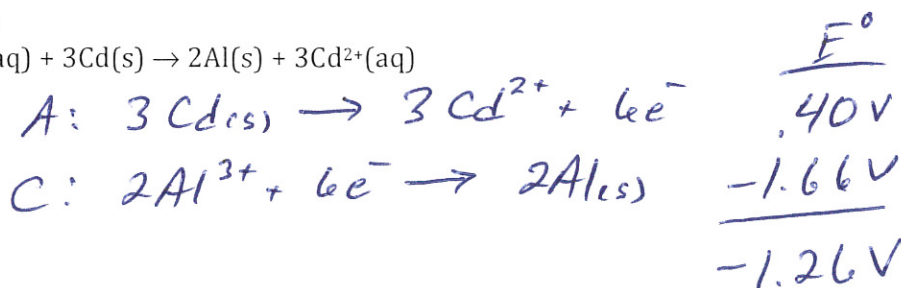
- a) +6.00 V d) +2.78 V
 b) -3.08 V e) -2.78 V
 c) +3.08 V



7. Calculate E° for the following reaction:



- a) -2.06 V d) -4.52 V
 b) +4.52 V e) -1.26 V
 c) +2.06 V



8. The standard cell potential for $3\text{Sn}^{4+}(\text{aq}) + 2\text{Al}(\text{s}) \rightarrow 3\text{Sn}^{2+}(\text{aq}) + 2\text{Al}^{3+}(\text{aq})$ is $E^\circ = 1.81 \text{ V}$.

What is E_{cell} when

$[\text{Sn}^{4+}] = 1.0$, $[\text{Sn}^{2+}] = 1.0 \times 10^{-2}$, and $[\text{Al}^{3+}] = 1.5 \times 10^{-3}$ at 298 K.

a) 1.70 V

d) 1.86 V

b) 1.76 V

☒ e) 1.93 V

c) 1.81 V

$$Q = \frac{[\text{Sn}^{2+}]^3 [\text{Al}^{3+}]^2}{[\text{Sn}^{4+}]^3} = \frac{[.01]^3 [0.0015]^2}{[1.0]^3} = 2.25 \times 10^{-12}$$

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0592}{n} \log Q$$

$$E_{\text{cell}} = 1.81 - \frac{0.0592}{6} \log 2.25 \times 10^{-12} = 1.925$$

9. Predict the product at the anode when electric current is passed through a solution of KI.

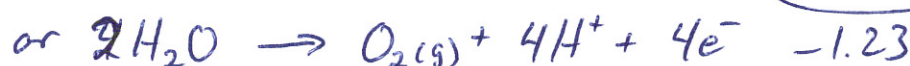
☒ a) $\text{I}_2(\text{l})$

d) $\text{K}(\text{s})$

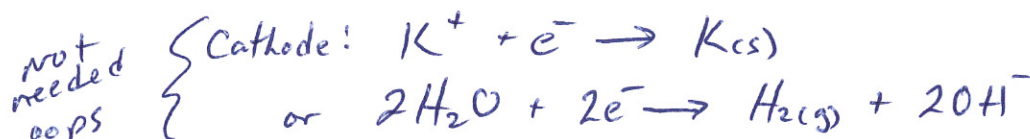
b) $\text{K}^+(\text{aq})$

e) $\text{O}_2(\text{g})$

c) $\text{H}_2(\text{g})$



$$\begin{array}{c} E^\circ \\ -1.535 \\ -1.23 \end{array}$$



10. If electric current is passed through aqueous LiBr, the product at the cathode would be _____ and the product at the anode would be _____.

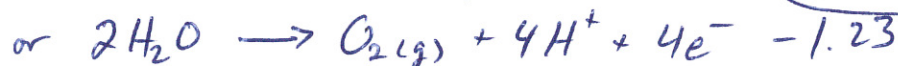
a) $\text{H}_2\text{O}(\text{l})$, $\text{Li}^+(\text{aq})$

d) $\text{Br}_2(\text{l})$, $\text{H}_2(\text{g})$

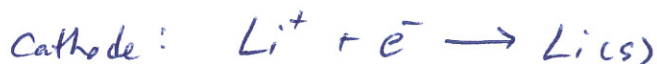
b) $\text{Br}_2(\text{l})$, $\text{Li}(\text{s})$

☒ e) $\text{H}_2(\text{g})$, $\text{Br}_2(\text{l})$

c) $\text{Li}(\text{s})$, $\text{Br}_2(\text{l})$



$$\begin{array}{c} E^\circ \\ -1.08 \\ -1.23 \end{array}$$



$$\begin{array}{c} -3.045 \\ -0.828 \end{array}$$

11. How long would it take to deposit 1.36 g of copper from an aqueous solution of copper(II) sulfate by passing a current of two amperes through the solution?

- a) 2070 sec d) 736 sec
b) 1.11×10^{-5} sec e) 1030 sec
c) 2570 sec

$$\frac{1.36 \text{ g Cu} \mid 1 \text{ mol Cu} \mid 2 \text{ mol } e^- \mid 96,500 \text{ C} \mid 1 \text{ s}}{63.55 \text{ g Cu} \mid 1 \text{ mol Cu} \mid 1 \text{ mol } e^- \mid 2 \text{ C}} = 2065 \text{ s}$$

$\swarrow \text{Cu}^{2+}$
 \uparrow
 2 amps

12. If a current of 6.0 amps is passed through a solution of Ag^+ for 1.5 hours, how many grams of silver are produced?

- a) 0.60 g d) 3.0 g
b) 36 g e) 1.0 g
c) 0.34 g

$$\frac{1.5 \text{ hr} \mid 60 \text{ min} \mid 60 \text{ s} \mid 6.0 \text{ C} \mid 1 \text{ mol } e^- \mid 1 \text{ mol Ag} \mid 107.87 \text{ g Ag}}{1 \text{ hr} \mid 1 \text{ min} \mid 1 \text{ s} \mid 96,500 \text{ C} \mid 1 \text{ mol } e^- \mid 1 \text{ mol Ag}} = 36.2 \text{ g Ag}$$

$\swarrow 6 \text{ amps}$ $\swarrow \text{Ag}^+$

13. $\text{Zn(s)} + \text{Cl}_2(\text{g}) \rightleftharpoons \text{Zn}^{2+}(\text{aq}) + 2\text{Cl}^-(\text{aq})$

An electrochemical cell based on this reaction has a cell voltage, E° , of 2.12 V. Which change could make the cell voltage greater than 2.12 V?

- a) add more Zn(s)
b) add more $\text{Cl}^-(\text{aq})$ ions
c) decrease the concentration of $\text{Zn}^{2+}(\text{aq})$ ions
d) decrease the partial pressure of Cl_2

$$Q = \frac{[\text{Zn}^{2+}][\text{Cl}^-]^2}{1}$$

$$E_{\text{cell}} = E_{\text{cell}}^\circ - \frac{0.0592}{n} \log Q$$

decreasing $[\text{Zn}^{2+}]$ decreases Q ,
which increases E_{cell} .